...turbance ...serious consequences, including death. ...ood or body can

...compression Sickness and Carbon Monoxide Poisoning

Disturbance	Effects on the Body	Equilibrium Before and After Treatment
...compression sickness and scuba diving	• As a diver descends, pressure increases in the diver's lungs as water pressure increases. • Solubility of gases in the blood increases as a result. • High concentrations of nitrogen affect the nervous system, causing confusion, drowsiness, and impaired judgment. • Increased dissolved oxygen due to increased solubility at deep temperatures can lead to seizures. • As the diver ascends, nitrogen leaves the blood as solubility decreases. Ascending too quickly means that nitrogen cannot be exhaled quickly enough. Nitrogen bubbles collect in joints and other parts of the body, damaging tissue and causing pain known as "the bends."	• $N_2(aq) \rightleftharpoons N_2(g)$ • Before treatment, equilibrium shifts to the right as a diver ascends, resulting in "the bends." • A hyperbaric chamber, which controls the concentration and pressure of a person's surroundings, can be used for treatment. • Equilibrium shifts back to the left as the pressure in the chamber increases. • Solubility of nitrogen gas increases in the blood and nitrogen bubbles dissolve. • Pressure is slowly decreased again to allow time for the diver to exhale gaseous nitrogen.
Carbon monoxide poisoning	• Carbon monoxide, CO(g), is a poisonous gas that is formed when incomplete combustion of a hydrocarbon occurs. • Structurally, carbon monoxide and oxygen are similar. • Carbon monoxide can bind to **hemoglobin,** which is the protein that carries oxygen in the blood. CO(g) blocks oxygen from binding to hemoglobin because it binds much more readily to hemoglobin than to oxygen—its equilibrium constant is 250 times greater than that of oxygen. Therefore, breathing even small amounts of CO(g) can be deadly. • Ultimately, carbon monoxide poisoning results in insufficient levels of oxygen being transported to body cells, resulting in cell death.	• The equilibria involved can be represented in simplified form, as shown: $O_2(g) + HbCO(aq) \rightleftharpoons O_2(g) + Hb(aq) + CO(aq) \rightleftharpoons HbO_2(g) + CO(g)$ • Equilibrium shifts to the left when a person is exposed to carbon monoxide as the carbon monoxide binds to hemoglobin. This binding reaction is strongly favoured. • In a hyperbaric chamber, high-pressure, high-oxygen conditions are created. • In these conditions, the equilibrium shifts to the right. Oxygen binds to hemoglobin and carbon monoxide gradually leaves the affected person's body.

Equilibrium Systems and Coral Reefs (7.4)

Coral reefs are rich ecosystems that support a diverse range of aquatic life. These reefs depend on equilibria in their natural environment that are affected by changes caused by human activities. The equilibria in seawater can be combined and summarized in simplified form as the following equation:

$$CO_2(g) + CO_3^{2-}(aq) + H_2O(\ell) \rightleftharpoons 2HCO_3^{-}(aq)$$

The burning of fossil fuels has led to increasing levels of carbon dioxide in the atmosphere. This has led to decreased concentrations of $CO_3^{2-}(aq)$ as equilibrium shifts to the right. The result of this has been a slowing of the growth of corals, because there is a smaller amount of carbonate ion available for coral polyps to make shells. (The overall equation for the shell-making process is shown below.)

$$Ca^{2+}(aq) + CO_3^{2-}(aq) \rightleftharpoons CaCO_3(s)$$

This process has already had devastating consequences for coral reef ecosystems worldwide and conservation projects are widespread.

Manufacturing Ammonia

The Haber-Bosch process for manufacturing ammonia was revolutionary because it was one of the first industrial processes that used principles of chemical equilibrium to influence product formation. Fritz Haber, a German chemist, deduced that reaction rate could be increased by heating a system. In the case of the synthesis of ammonia, though, the reaction is exothermic. This should result in an ever-increasing equilibrium shift favouring reactant formation as more thermal energy is generated. Haber recognized this potential issue. He determined that by increasing the pressure of the system, the shift toward reactant formation could be overcome, as predicted by Le Châtelier's principle.

To increase reaction rate and yield of ammonia, Haber experimented to find a suitable catalyst to speed up the reaction. He also designed a way to remove ammonia from the system, thus continually shifting the equilibrium to the right. These adjustments to the reaction conditions result in a significant increase in ammonia yield.

System Changes Used in the Haber-Bosch Process

Overall Reaction: $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + 92$ kJ		
Change	**System Shift**	**Explanation**
Pressure ⇑	To the right	As the pressure of a system increases, the equilibrium system will shift to *decrease the number of gas molecules*. This results in fewer collisions, which counteracts changes in pressure.
$NH_3(g)$ ⇓	To the right	As ammonia is continually removed from the system, the equilibrium is continually shifted to the right, favouring the production of ammonia.
Addition of catalyst	No shift	Adding a catalyst increases reaction rate in both directions, which results in an increased rate of production of ammonia due to the other factors shifting the reaction to the right.

The following diagram represents how the Haber-Bosch process works in practice.